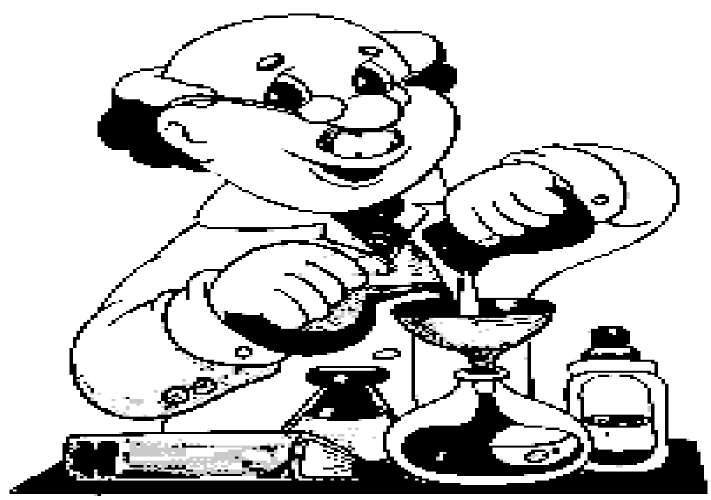
Geel 2000 Language Schools

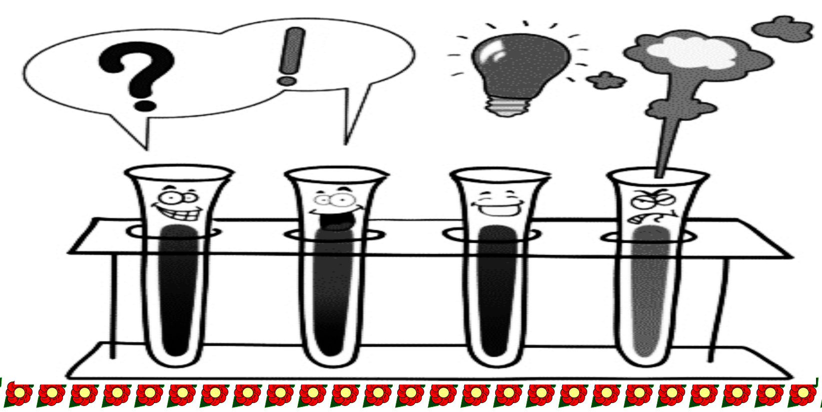
Science Department

Chemistry

First secondary

2023/2024





Contents

Unit one

Chemistry is the central science

Unit One

Chapter (1): Chemistry and measurement

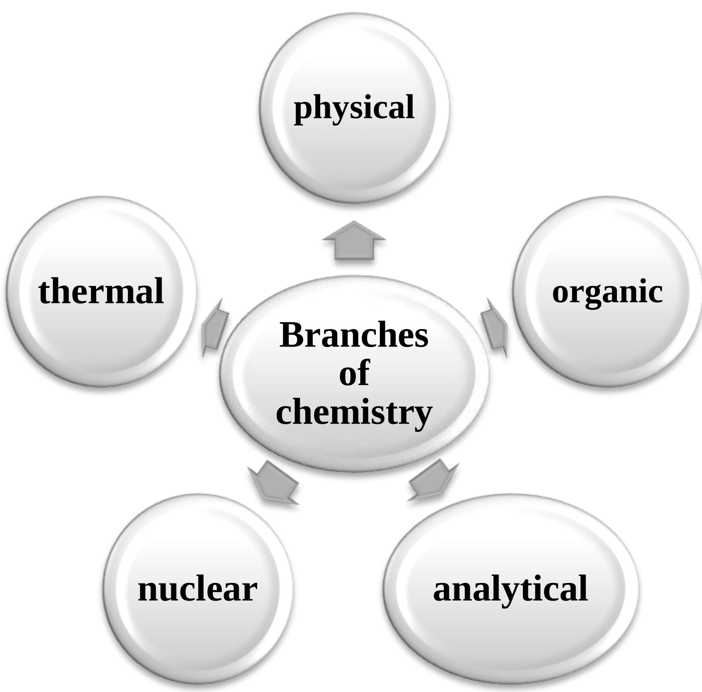
Science:

An organized structure of knowledge that includes facts, principles, laws and scientific theories.

Chemistry:

Science that studies the composition, properties of matter, changes that occur to it and reactions between substances.

Branches of chemistry:



Fields of chemistry:

In ancient times

1-Metal and mining

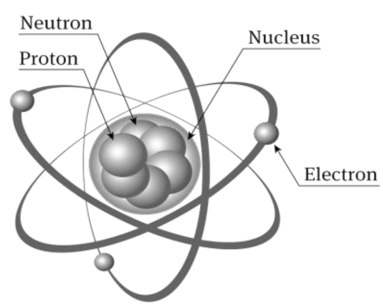
2-production of colors and glasses.

3-Tanning and dying clothes

4-Medicines.

5-Mummifying

Now days



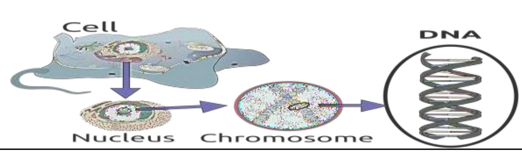
Study

1-properties of substances.

2-Structure of atoms and molecules.

3-Solving some problems as pollution

*Relation between chemistry and other branches of science:

1-Biology	**Biology:** Study the living organisms. **Chemistry:** Study reactions inside the body as digestion **Chemistry + biology = Biochemistry:** **Biochemistry:** Science study the chemical structure of proteins, fats and carbohydrate 

2-Physics	**Physics:** study natural phenomena as motion , force, light. **Chemistry + physics = Physical chemistry** Study -Properties of substances. -Structure of these substances. 
3-Medicine and pharmacy	**Medicine:** they are chemical substances that have healing properties. **Chemistry studies:** Nature and function of hormones and enzymes in the body. 
4-Agriculture	**Chemistry helps in** 1-Choosing the suitable soil. 2-Suitable fertilizer. 
5-Future	Discover substances with extraordinary properties through nano chemistry

Measurement in chemistry

Measurement:

Comparison of unknown quantity with another known one.

*Results of measurement:

1-Numerical value: to describe the physical quantity

2-Measuring unit

Measuring unit:

It is a certain portion of a certain physical quantity

*Importance of measurement:

1-Gaining information about substances.

2-Monitoring and protection.

3-Evaluate a situation and suggest medicine as (glucose in blood)

Measuring tools

1-Sensitive balance:

Use: measure the mass of substances

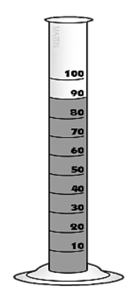
Types: Digital balances

Most common: Top loading balances

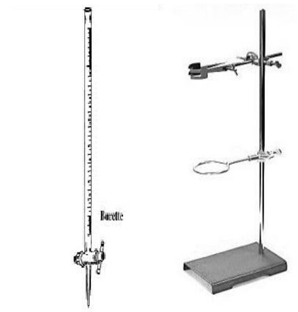
2-Graduated cylinder:

Use: 1-Measure the volume of liquids with high accuracy

2-measure the volume of solid bodies.



3-Burette:

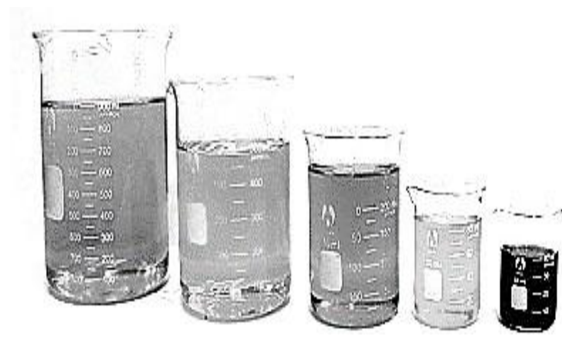


Description: Long glass tube with two opening the graduation zero is close to upper opening

Use: Titration

Note: It should be fixed on a holder with a metallic Base

4-Beakers:



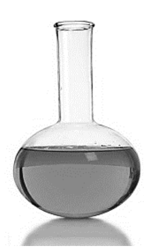
Description: transparent beakers made of pyrex glass

Use: 1-Measure approximate volume of solutions

2-Transporting solutions

5-Flasks:

Description: has many shapes

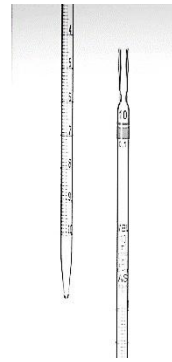
Type	Conical flask	Round-bottom flask	Volumetric flask
Use	Titration 	Preparation and distillation 	Prepare solution with accurate concentration 

6-Pipette:

Description: Long glass tube opened from the two sides

Use: Measure and transport certain volume of solution

Note: vacuum should be used to avoid harms



7-Tools for measuring PH:

PH:

-It is the measurement that determine if the substance is acid or base or neutral.

-It is the measurement of concentration of hydrogen ions in solution.

-Tools to detect PH:

1-litmus paper   (by changing their color)

2-Digital apparatus (more accurate and it measures PH directly)

-PH < 7   acid

-PH = 7   neutral

-PH > 7   base

PH

0   acid   7   basic   14

- **PH meter is more accurate in measuring PH of a solution. (G.R)**

Because PH meter can determine if the solution is acid or base and also determine the concentration of hydrogen ions in the solution while PH tape is used to know if the substance is acid or base only

Chapter (2): Nanotechnology and chemistry

Nanotechnology

Nano-------------derives from Greek word **Nanos** and means dwarf or very small

Technology---------------applied application of knowledge in certain field.

Nanotechnology:

It is the technology of very small substances and it specialized in treating the substance on Nano measure to produce new, useful, and unique properties.

The Nano is a unique measuring unit:

1 milli = $1x10^{-3}$m

1 micro = $1x10^{-6}$m

1Nano = $1x10^{-9}$m (1 Nano = one part of a billion part of meter)

Why the Nano scale is unique in measurement?

1-The properties of substance as (color, transparency, ability to conduct heat and electricity

2-Speed of chemical reaction, toughness, elasticity,….) change completely in Nano scale.

3-The substance gain new and unique properties.(prop. Change with changing Nano volume).

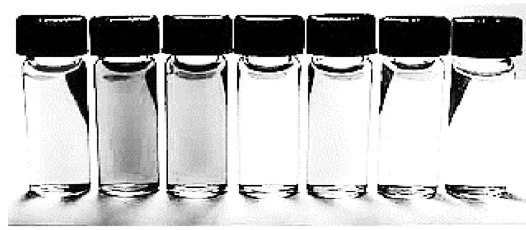
3-Nano substances can be used in new and uncommon applications.

Critical Nano volume:

The volume in which the unique Nano properties of the substance appear and is located between (1-100 nm).

- So the properties of substances in Nano scale is **volume dependent properties.**

- **Examples on substances in Nano scale:**



1-Nano gold:

- The gold is **yellow** in color and bright in normal scale

- Nano gold takes **different colors** according to their Nano volume (It may be red, green , orange and blue).

Because the reaction of Nano gold with light is different from reaction of gold in macro volume.

2- Nano copper:

The **hardness** of nano copper is more than its hardness in macro measurement

The speed of reaction in Nano scale:

In the Nano volume of the substance, the ratio **increases** between the surface area to volume so the number of atoms exposed to reactions increases so the speed increase and the substances gain new properties.

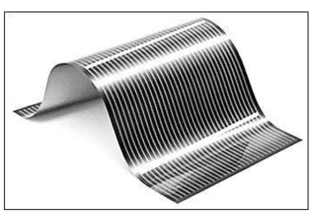
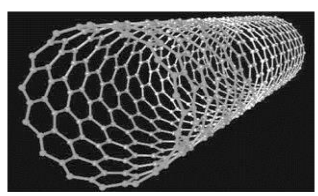
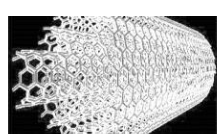
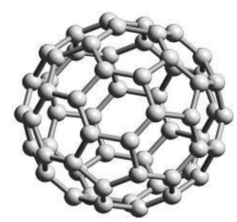
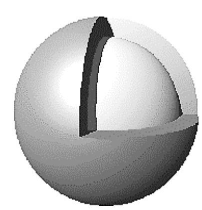
- When substance changes from macro measurement to nano measurement **surface area increases** while **volume remains constant**

Nanochemistry

Nanochemistry:

It is the branch of Nano science , it deals with chemical applications Of Nano substances.

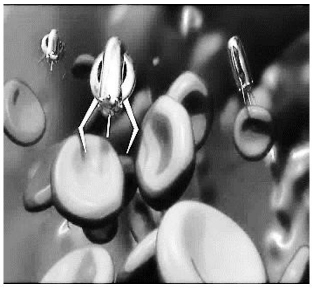
***Nano substances can be classified according to the dimensions into**

Type	**1)One dimensional Nano substances**	**2)Two dimensional Nano substances**	**3)Three dimensional Nano substances**
Definition	They are nano substances with one Nano dimensionless than 100 nm.	They are Nano substances with two dimensions each of them less than 100 nm.	They are nano substances with three dimensions each of them less Than 100 nm.
Examples	-Thin films  -Nano wires and fibers 	-carbon Nanotube  -Multi carbon nano tube 	-Bucky ball C_{60}  -Nano shell 

Uses	***Thin films:** are used in **1-**Painting surfaces **to protect them from rust.** 2-Packing food Products **to protect them from getting spoiled or rotting.** ***Nano wires:** are used in electrical circuits. ***Nano fibers:** are used in production of water filters	***Carbon Nano tubes:** Are 1-Good conductors of electricity than copper. 2-Good conductors of heat than diamond 3-Stronger and lighter than steel due to powerful bond between its molecules. 4-Connected easily to protein so they can be used in making biological sensor devices which are sensitive to certain molecules.	***Bucky ball C60** Used as 1-Carrier for medicine in the body. **Due to its hollow structure It can match With a molecule of medicine** but its outer part resist the reaction of the medicine with other molecules in the body.

Applications on Nano technology

1) Medical field:



1-The early diagnosing of diseases and picturing organs and tissues.

2-Deliver medicine to the infected tissues and cells which increases the chances of healing and reduce harmful effects.

3-Producing very minute devices for dialysis that can implanted in the body.

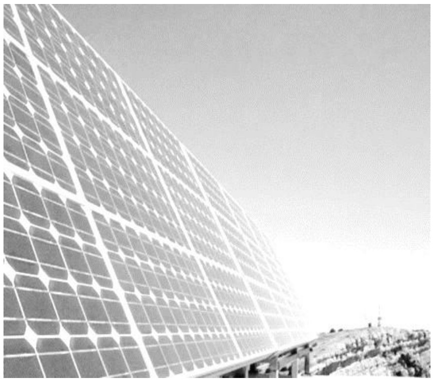
4-Producing Nano robots that sent into blood streams and remove blood clots from veins without surgical interferences.

2) Agricultural field:

1-Identify bacteria in nutrients and preserving food.

2-Improve nutrients , pesticides and medicines for plants and animals.

3) Energy field:



1-Produce solar cells using Nano silicon that has high ability to transform energy without leakage of heat energy.

2-Producing Hydrogen fuel cells that are low in cost and high on performance.

4) Industrial fields:

1-Producing invisible Nano molecules that acquire glass and ceramic property of self cleaning.

2-Producing Nano substances to purify ultraviolet rays in order to improve sun block cosmetics and creams.

3-Producing a Nano wrapping technology in the form of paints and sprays that work to form layers of coverings that protect the screens of electrical devices from scratching.

4-Producing repellent tissues for stains and distinguished with self-cleaning.

4) Communications field:

1-Producing wireless Nano devices, mobiles, and satellites.

2-Decreasing the size of the transistors.

3-Producing electric chips that are distinguished with a high storage capability.

5) Environmental fields:

1-Producing Nano filters that work on purifying the air, water, solving the problem of nuclear wastes and removing the dangerous elements from industrial wastes.

Harmful effects of nanotechnology

1) Medical effects:

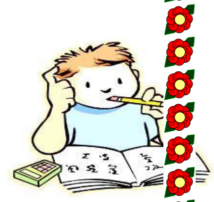
Very minute nano particles may enter the body of human or animal through cell membrane of skin or lungs causing diseases.

2) Environmental effects:

During production of Nano substances some wastes may be suspended in the air, water and soil causing pollution.

3) Social effects:

It may cause social inequality between rich countries and developed countries.



Unit two
Quantitative chemistry

Unit two

Chapter (1): The mole and the chemical equation

Part (1): Ionic equations

The following table shows the valency of some metals

Element Monovalent Valency (+1)	Element Divalent Valency(+2)	Element Trivalent Valency (+3)
Lithium (Li) Sodium (Na) Silver (Ag) Potassium (K)	Mercury (Hg) Magnesium (Mg) Calcium (Ca) Lead (Pb) Iron (Fe) Copper (Cu)	Aluminum (Al) Iron (Fe) Gold (Au)

Table of atomic groups

Atomic group	Symbol	Valency	Atomic Group	Symbol	Valency
Hydroxide	OH^-	-1	Sulphate	SO_4^{-2}	-2
Nitrate	NO_3^-	-1	Carbonate	CO_3^{-2}	-2
Nitrite	NO_2^-	-1			
Bicarbonate	HCO_3^-	-1	Phosphate	PO_4^{-3}	-3
Ammonium	NH_4^+	+1			

Chemical equation:

A group of chemical symbols and formulas of the reactants and products. They connected by an arrow between them that express the direction of this reaction and carry the reaction condition.

$$2Mg_{(s)} + O_{2\,(g)} \xrightarrow{\triangle} 2MgO_{(s)}$$

- The equation includes the physical states written at the bottom left of the chemical symbols.

Solid	S
Liquid	L
Gas	G
Aqueous solution	Aq

- **The equation must be balanced. (G.R)**

To achieve the law of mass conservation.

Example: Try to balance these reactions

1) $H_2 + O_2 \longrightarrow H_2O$
2) $N_2 + H_2 \longrightarrow NH_3$

Answer:

1) $2H_2 + O_2 \longrightarrow 2H_2O$
2) $N_2 + 3H_2 \longrightarrow 2NH_3$

Ionic equations

It is the chemical equation in which reactants and products are written in the form of ions.

1) Dissolving equations:

As dissolving sodium chloride in water

$NaCl_{(s)} \xrightarrow{water} Na^+ + Cl^-$

2) Neutralization reaction:

It is the reaction between acid and base to produce salt and water.

Example:

Reaction between sulphuric acid and sodium hydroxide.

$H_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

$2H^+_{(aq)} + \cancel{SO_4}^{2-}_{(aq)} + \cancel{2Na}^+_{(aq)} + 2OH^-_{(aq)} \longrightarrow \cancel{2Na}^+_{(aq)} + \cancel{SO_4}^{-2}_{(aq)} + 2H_2O_{(l)}$

$2H^+_{(aq)} + 2OH^-_{(aq)} \longrightarrow 2H_2O_{(l)}$

3)The ionic equation for precipitation reaction:

As precipitation of silver dichromate on adding potassium dichromate solution to silver nitrates solution.

$$K_2Cr_2O_{7(aq)} + 2AgNO_{3(aq)} \longrightarrow 2KNO_{3(aq)} + Ag_2Cr_2O_{7(S)}$$

$$\cancel{2K^+_{(aq)}} + Cr_2O_7{}^{2-}{}_{(aq)} + 2Ag^+_{(aq)} + \cancel{2NO_3^-{}_{(aq)}} \longrightarrow \cancel{2K^+_{(aq)}} + \cancel{2NO_3^-{}_{(aq)}} + Ag_2Cr_2O_{7(S)}$$

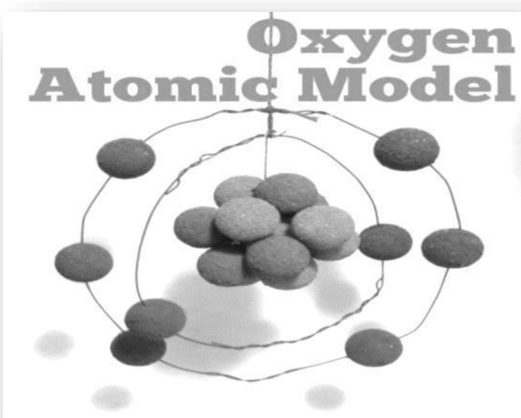
$$2Ag^+_{(aq)} + Cr_2O_7{}^{2-}{}_{(aq)} \longrightarrow Ag_2Cr_2O_{7(S)}$$

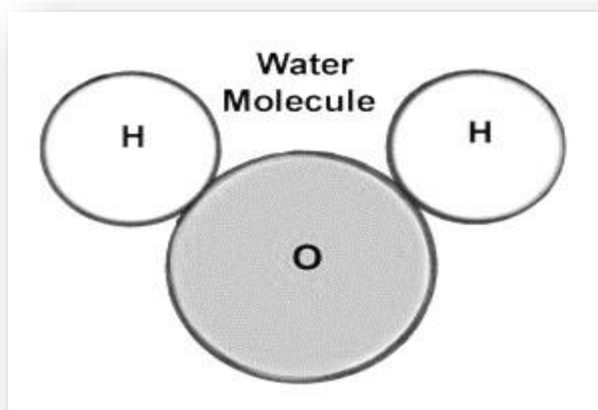
Molecule:

It is the smallest part of the substance that can be found in a single form and carry the properties of matter.

Atom:

It is the smallest building unit of the substance that can participate in chemical reactions





Unit two

Chapter (1): The mole and the chemical equation

Part (2): The mole and molar mass

The mole:

It is the amount of substance that contains Avogadro number.

- If the substance is in the form of **atoms** , the mass of one atom is called **atomic mass**

It is very small and measured by **atomic mass unit (a.m.u.)**.

- If the atomic mass of carbon atom (C) = 12 a.m.u., then **one mole** of carbon

atom = 12 **grams** of carbon atoms.

- If the substance is in the form of **molecules** , then the mass of one molecule is called **molecular mass .** It is equal the sum of atomic mass of atoms forming this molecule.

Molecular mass:

It is the sum of the atomic mass of the atoms forming the molecule.

Example:

Calculate the molecular mass of carbon dioxide (CO_2).If you know that the atomic mass of oxygen is 16 and carbon is 12.

Answer:

Molecular mass of CO_2 = (atomic mass of carbon) + (2 x atomic mass of oxygen)

$= (12) + (2 \times 16)$

$= (12) + (32) = 44$ a.m.u.

One mole of CO_2 = 44 g

- If we use 44 g carbon dioxide, this means that you use one mole of it.

- If we use 22 g carbon dioxide, this means that you use half mole of it. And so on
- In ionic compounds the building units can expressed in **formula unit** not molecules. So ionic compounds have **formula unit mass** not **molecular mass.**

Example: **Calculate** the formula unit mass for ionic calcium chloride ($CaCl_2$).
If you know that the atomic mass of calcium ion is 40 and chloride ion is 35.5 .

Answer:

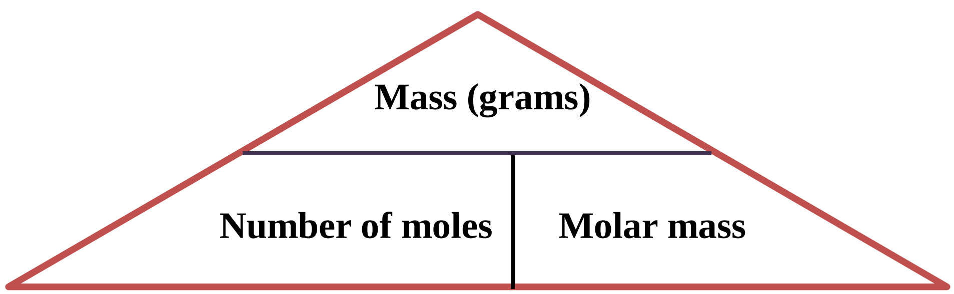
mass of $CaCl_2$ = (mass of calcium ion) + (2 x mass of chloride ion)

= (40) + (2 × 35.5)

= (40) + (71) = 111 a.m.u.

One mole of CO_2 = 111 g

$$\boldsymbol{Number\ of\ moles} = \frac{\textbf{mass of substance (gram)}}{\textbf{Mass of one mole of this substance (g/mol)}}$$



*** The mass of a mole (molar mass) different from one matter to another.(G.R)**
Due to the difference in composition from one matter to another.

- The mole of molecules of monatomic element is different from the mole of the same element if it is diatomic.

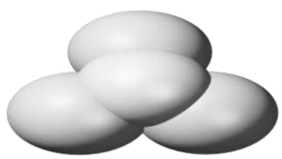
*The molar mass of oxygen molecules = 16 × 2 = 32 g

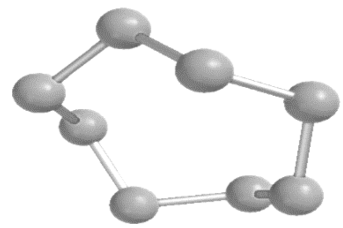
*The molar mass of oxygen atom = 16 × 1 = 16 g

- There are elements with different molecular composition due to difference in their physical state as

- phosphorus in vapour state formed from **four** phosphorus atoms (**P_4**), while in solid state it consists of one atom

-Sulphur in vapour state formed from **eight** sulphur atoms (**S_8**), while in solid state it consists of one atom





Calculation of the mass of reactants and products:

Example:

Calculate the mass of magnesium needed to react with excess amount of oxygen to produce 160g of magnesium oxide. [Mg = 24, O=16]

Answer:

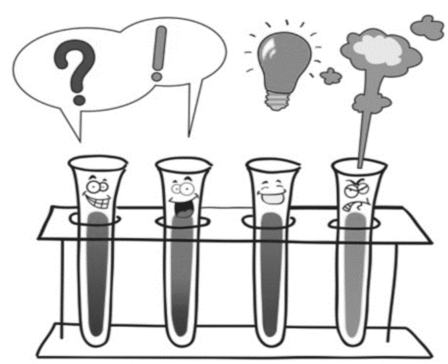
$2Mg + O_2 \longrightarrow 2MgO$

2moles   2 moles

2x24=48g   2(24+16)=80g

???   160g

Mass of magnesium= $\frac{160 \times 48}{80} = 96\ g$



Unit two

Chapter (1): The mole and the chemical equation

Part (3): The mole and Avogadro's number and volume of gases

The mole and Avogadro's number:

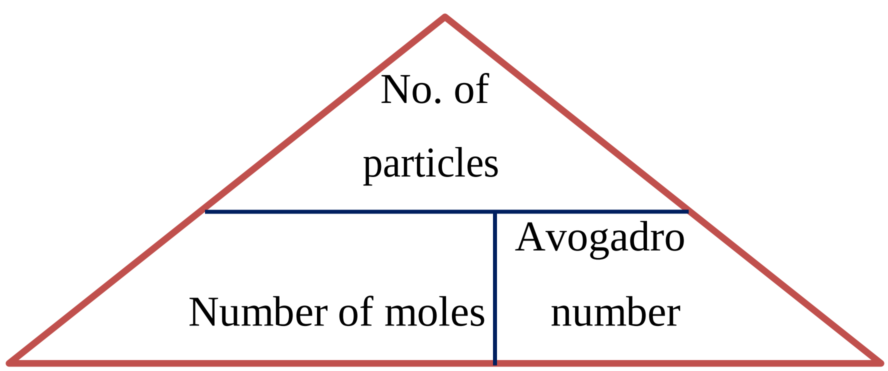
Avogadro reached that the number of atoms, molecules or ions found in one mole is a constant number whatever the form of substance.

Avogadro number = 6.02×10^{23}

Avogadro number:

It is the number of atoms , molecules or ions found in one mole of the substance and equals 6.02×10^{23} (atoms, molecules or ions).

$$\textit{Number of mole} = \frac{\textbf{number of particles}}{\textbf{Avogadro number}}$$



Exercise (1):

Calculate the number of molecules of 2 mol of CO_2 gas.

Answer:

Number of molecules = number of moles × Avogadro number

$= 2 \times 6.02 \times 10^{23} = 12.04 \times 10^{23}$ molecules

---

Exercise (2):

Calculate the number of carbon atoms found in 50 g of calcium carbonates

[Ca = 40, C = 12,O = 16]

Answer:

1 mole of calcium carbonate $CaCO_3 = 40 + 12 + (16 \times 3) = 100$ g

1 mole $CaCO_3 \xrightarrow{\text{contains}}$ 1 mol of carbon atoms

100 g $CaCO_3 \xrightarrow{\text{contains}}$ (6.02×10^{23}) atoms

50 g $CaCO_3 \xrightarrow{\text{contains}}$ × carbon atoms

$$x = \frac{(6.02\ x\ 10^{\wedge}23\ x\ 50)}{100} = 3.01 \times 10^{23} \text{ atom}$$

Exercise (3):

Calculate the number of magnesium oxide molecules produced from reaction of 24 grams of magnesium with excess amount of oxygen. [Mg = 24]

Answer:

$2Mg + O_2 \longrightarrow 2MgO$

2moles $\longrightarrow$ 2 moles

2 x 24= 48g $\longrightarrow$ $2x\ 6.02\ x\ 10^{23}$

24 g $\longrightarrow$??

Number of magnesium oxide molecules= $\frac{24\ x\ 2\ x\ 6.02x10^{\wedge}23}{48} = 6.02 \times 10^{23}$ molecules.

The mole and the volume of gas

- Solid or liquid matter has a definite volume.
- The volume of gas equal the volume of the container it occupies
- The mole of any gas in standard temperature and pressure (STP) occupies a certain volume =22.4 liters.

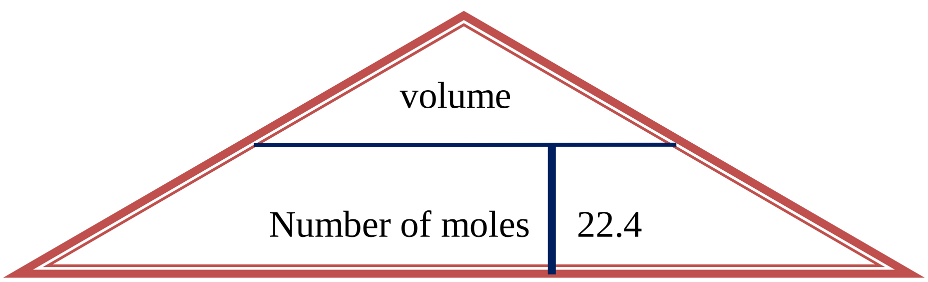
- **(STP) means**

1-Temperature equals 273 k or 0^{o} c

2-pressure = 760 mmHg (normal atmospheric pressure = 1 atm.p)

3-Concentration = 1 molar

$$\textit{Number of moles} = \frac{\textbf{volume}}{\textbf{22.4 L}}$$



Example(1):

Calculate the volume of 3 moles of oxygen gas

Answer:

Volume = number of moles × 22.4 = 3 × 22.4 = 67.2 L

Example(2): **Calculate the volume of oxygen needed to produce 90 g of water by reacting with an excess amount of hydrogen at the standard temperature and pressure (STP) [H = 1 , O = 16]**

Answer:

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O(l)$$

$$1\text{ mol} \longrightarrow 2\text{ mol}$$

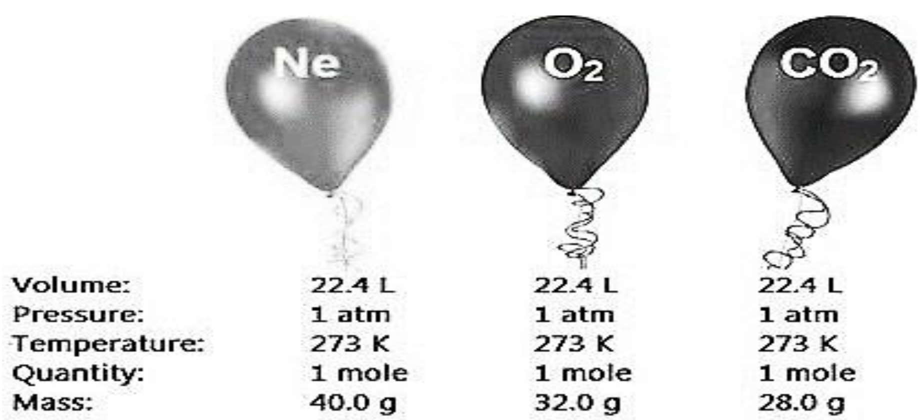
$$22.4L \longrightarrow 2[(2 \times 1) + 16] = 36\text{ g}$$

$$?? \longrightarrow 90\text{ g}$$

Volume of oxygen $= \frac{22.4 \times 90}{36} = 56$ L

Avogadro Hypothesis:

Equal volume of different gases contain the same number of molecules under the same standard temperature and pressure (STP).



Avogadro Law:

At constant temperature and pressure the volume of gas is directly proportional to its number of moles

Limiting reactant

Limiting reactant

It is the reactant which is completely consumed during chemical reaction

Example:

When magnesium reacts with oxygen according to the equation

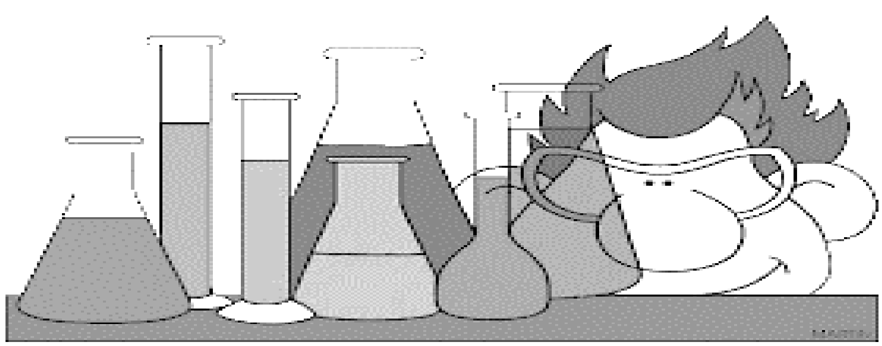
$2Mg + O_2 \longrightarrow 2MgO$

What is limiting reactant when 32 g of oxygen reacts with 12 g of magnesium?

[Mg = 24 , O=16]

$2Mg \longrightarrow$	$2MgO$	$O_2 \longrightarrow$	$2MgO$
2moles	2moles	1 moles	2 moles
2×24=48g	2(24+16)=80g	32 g	80 g
12g	??	32 g	??
Mass of magnesium oxide=20 g		mass of magnesium oxide = 80 g	

Limiting reactant is magnesium



Unit two

Chapter (2): Calculation of chemical formula

Part (1): Calculation of chemical formula

Weight percentage:

It used to calculate the ratio of each component from the component of certain sample.

$$\textit{Matter weight percentage} = \frac{\textit{matter mass in the sample}}{\textit{total mass of the sample}} \times 100$$

Example:

Calculate the weight percentage of nitrogen in one mole of ammonium nitrates fertilizers [N = 14 , H=1 , O=16]

Answer:

Molar mass of ammonium nitrate NH_4NO_3 = (14 + (1 ×4) +14 + (3 × 16))= 80 g

Each one mole of ammonium nitrate contains 2 mol nitrogen = (2 × 14) = 28g

$$\textit{weight percentage of nitrogen} = \frac{28}{80} \times 100 = 35\%$$

By calculating the weight percentage of oxygen and hydrogen.

Weight percentage of oxygen = 60%

Weight percentage of hydrogen = 5%

- The sum of the weight percentage of a compound must be equal 100

Example:

Calculate the mass of iron found in one ton (1000 kg) of hematite ore Fe_2O_3 , if you know that the weight percentage of iron is 58%

Answer:

58% means that each

100 ton ore $\xrightarrow{\text{contains}}$ 58 ton iron

1 ton $\xrightarrow{\text{contains}}$??? ton

$$x = \frac{1 \times 58}{100} = 0.58 \text{ ton} = 580 \text{ kg}$$

Example:

Calculate the number of moles of carbon in an organic compound containing only carbon and hydrogen. If you knew that the weight percentage of carbon in this compound is 85.71% and the molar mass of this compound is 28 g (C=12 , H= 1).

Answer:

There is 85.71 g carbon -------------- in 100 g of the sample

So there is × g carbon ------------------ in 28g

X = (28 × 85.71) / 100 = 24 g

Number of carbon moles = 24/12 = 2 mol

Another answer:

$$carbon\ mass = \frac{carbon\ weight\ percentage\ x\ molar\ mass\ of\ the\ compound}{100}$$

$$= \frac{85.71\ x\ 28}{100} = 24 \text{ g} = 2 \text{ mol}$$

Calculation of chemical formula

Empirical formula:

A formula expressing the simplest ratio of true numbers between the atoms of elements which formed the compound.

Example:

The molecular formula of propylene is C_3H_6 ----------- That means that the molecule of propylene is formed of 6 atoms of hydrogen and 3 atoms of carbon with ratio of 6(H) : 3(C).

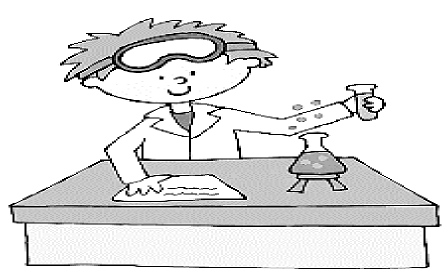
By simplifying this ratio to its true value the ratio becomes

2 (H) : 1 (C) so the empirical formula is CH_2

- Sometimes the empirical formula is similar to chemical formula like (carbon dioxide CO_2) – (Nitric oxide NO)

- The empirical formula of two different compounds may be the same like acetylene C_2H_2 and benzene C_6H_6 . Both of them has empirical formula (CH)

- The empirical formula of the compound can be calculated in terms of weight percentage of elements that represent that mass of elements found in 100 g



Example:

Calculate the empirical formula of a compound containing nitrogen with a weight percentage of 25.9 % and oxygen with a weight percentage of 74.1 % knowing that (N = 14 , O = 16)

Answer:

	N	:	O
Number of moles	$\frac{25.9}{14}$	:	$\frac{74.1}{16}$
	$\frac{1.85}{1.85}$	:	$\frac{4.63}{1.85}$
	1	:	2.5 *(x 2)*

The empirical formula is N_2O_5

Molecular formula:
Is a symbolic formula of the molecule of the element, or molecule or formula unit. It express the actual type and number of atoms or ions that form this molecule or unit.

$$\textit{Number of units of the emperical formula} = \frac{\textit{molar mass of the compound}}{\textit{molar mass of the emperical formula}}$$

Example:

Chemical analysis of acetic acid prove that it is formed from 40% carbon, 6.67% hydrogen , and 53.33% oxygen. If the molecular molar mass of it is 60 g find the molecular formula of the acid knowing that (C= 12, H=1, O=16)

Answer:

	C	H	O
Number of moles =	$\frac{40}{12}$	$\frac{6.67}{1}$	$\frac{53.33}{16}$
	3.33	6.67	3.33

Divide on the smallest number of moles

Ratio = 1 : 2 : 1

Empirical formula is CH_2O

Molecular mass of empirical formula= 16 +(1 x2) + 12 = 30 g

Number of units of empirical formula = $\frac{60}{30} = 2$

Molecular formula is $CH_2O \times 2 = C_2H_4O_2$

Unit two

Chapter (2): Calculation of chemical formula

Part (2): Practical and theoretical yield

Practical product and theoretical product

When we make a chemical reaction to obtain a certain substance the theoretical results that expected to get from the reaction is different from the produced substance practically (practically yield).

Theoretical yield:

It is the quantity of product calculated according to the chemical equation.

Practical yield:

It is the quantity of product that is actually produced from the reaction.

- **The practically yield usually less than the calculated amount theoretically.**

Due to

1-The product substances may evaporates.

2-Some of the product may clink on to the walls of of the reaction cylinder.

3-There are some side reactions (competitive reactions) that consume the product

4-The used substance may be not pure enough.

$$\textbf{Percentage of actual yield} = \frac{\text{practical yeild}}{Theoritical\ yeild} \times 100$$

Example:

Methyl alcohol is produced under high pressure through the following reaction

$$CO_{(g)} + 2H_{2\,(g)} \xrightarrow{\triangle} CH_3OH_{(l)}$$

If 6.1 g of methyl alcohol is produced from a reaction of 1.2 g of hydrogen with abundance of carbon oxide, calculate the percentage of the actual yield (C= 12, O = 16, H=1)

Answer:

Molecular mass of $CH_3OH = 1\times4 + 16 + 12 = 32$ g

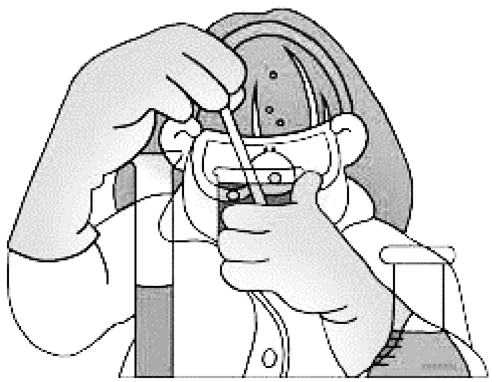
2mol of $H_2 \xrightarrow{\text{produce}}$ 1 mol of CH_3OH

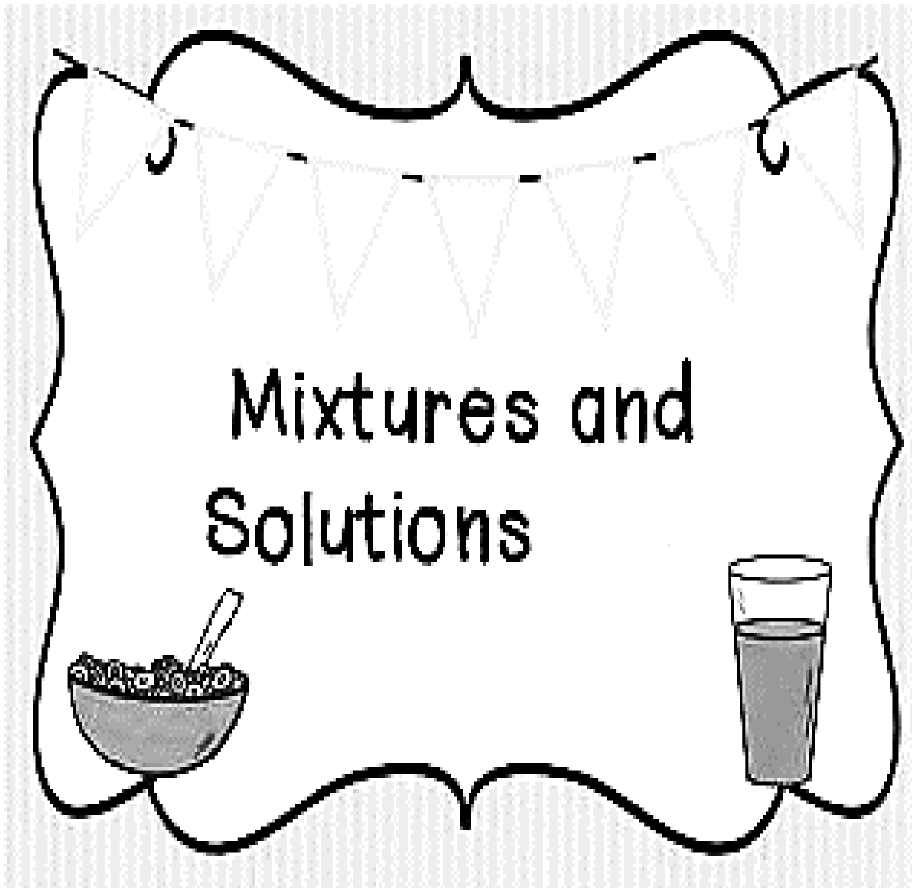
4g $\longrightarrow$ 32 g

1.2 $\longrightarrow$ x g

$X = \frac{32 \times 1.2}{4} = 9.6$ g

Percentage of yield $= \frac{6.1}{9.6} \times 100 = 63.54\ \%$





Unit Three

Solutions, Acids and bases

Unit three

Chapter (1): Solution and colloids

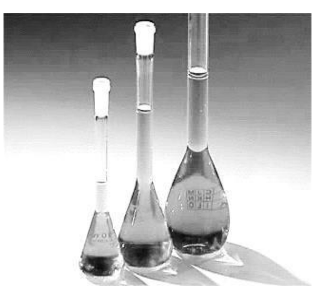
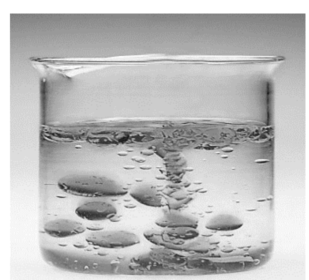
Part (1): Solutions

Types of mixtures

Homogenous mixtures **Heterogeneous mixture**

(solutions)

Colloids **Suspension**

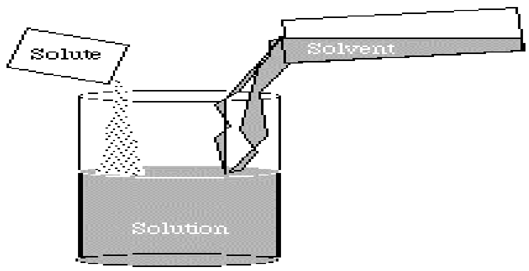
Solutions	Colloids	Suspension
-They are homogenous mixtures in which you cannot distinguished its components by naked eye or by electronic microscope **Examples:** -Table salt solution in water -Sugar in water -cobalt (II)chloride in water	-They are heterogeneous mixtures that carry the properties of solution and suspension -Components can be distinguished by microscope. **Examples:** -milk -blood -aerosols -hair gel -mayonnaise emulsion	-They are heterogeneous solutions in which you can distinguish its components by your eye **Examples:** -Table salt in kerosene -sugar in kerosene -cobalt (II)chloride in kerosene
		

Solutions

Solute + **Solvent** ⟶ **Solution**

(Small amount) **(Large amount)**

(Minor component) **(Major component)**

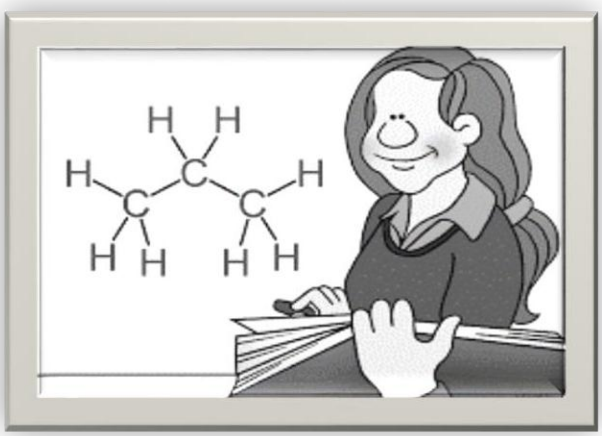


Solution

-It is a homogenous mixture of two or more substances.

***The sweet taste of sugar in sugary solution in all of its parts.(G.R)**

Because it is homogenous mixture all its parts contain the same amount of sugar.



Classifications of solutions

According to

1-The physical state of solvent

-Some times the word solutions is connected with a liquid state of the substance. But solutions may be in gas or liquid or solid state.

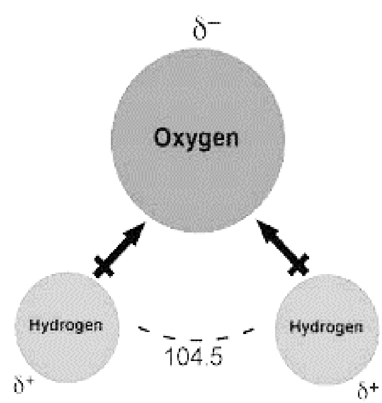
Type of Solution	Solute state	Solvent state	Examples
Gas	Gas	Gas	Air – natural gas
Liquids	Gas	Liquid	Soft drinks – oxygen dissolved in water
	Liquid		Alcohol in water Ethylene glycol(anti-freeze)in water
	Solid		Sugar or salt in water
Solid	Gas	Solid	Hydrogen in platinum or palladium
	Liquid		Silver amalgam $Ag_{(s)}$ / $Hg_{(l)}$
	Solid		Alloy of nickel – chrome alloy

Water is a polar solvent:

Water is a polar solvent because the electro negativity of oxygen is higher than hydrogen.

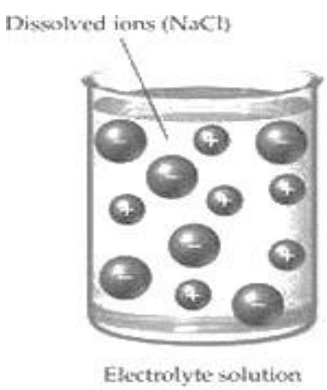
So oxygen carries a partial negative charge while hydrogen carries a partial positive charge

With angle 104.5^0 between them

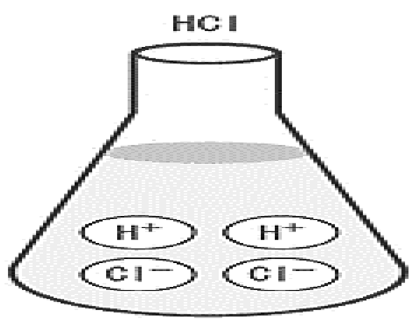
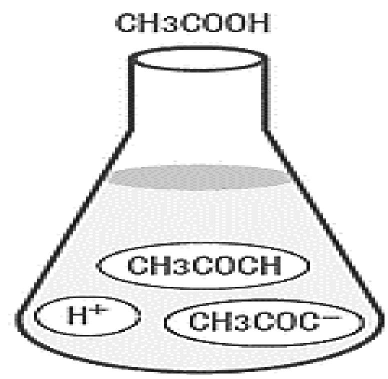


2-Ability to conduct electricity

- Solutions are classified according to conduction of electricity into

1-Electrolyte	2-Non electrolyte
The substance in which it solutions or its molten state conduct the electric current by the free ion movement. as (table salt solution)  	The substance in which its solution orits molten state do not conduct electricity because it doesn't have Freeions. as (Sugar solutions – ethyl alcohol) 

- **Electrolytes are classified into:**

Strong electrolytes	Weak electrolyte
They have the ability to conduct electricity to large extent as it is completely ionized (all its molecules are dissociated into ions) **Examples:** **1-Ionic compounds:** as -sodium chloride NaCl - sodium hydroxide NaOH **2-Polar covalent compounds:** as -Hydrogen chloride solution but hydrogen chloride in gas state doesn't conduct electricity.	Conduct the electricity to weak extent as it is partially ionized (Small parts of its molecules are dissociated into ions) **Examples:** **1-Ionic compounds:** as -Ammonium hydroxide NH_4OH **2-Polar covalent compounds:** as -Acetic acid CH_3COOH
 An example of strong electrolytes	 An example of weak electrolytes
$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$	$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$

3-Degree of saturation:

Classification of solutions according to degree of solubility

Unsaturated solutions	Saturated solutions	Super saturated solutions
It is the solution at which the solvent accepts more solute at a certain temperature.	It is the solution in which the solvent contains maximum amount of the solute at certain temperature.	It is the solution that accepts more of the solute after reaching saturation by heating

How can you prepare a saturated solution from a supersaturated solution?

a-Cooling

Cool the saturated solution and leave it for a short time, the excess solute will be precipitated.

b-Crystallization

Put small crystals from the solute in the supersaturated solution and leave it for a short time, the solute molecules will precipitates as crystals on the surface of seeding crystals.

Dissolving process

Dissolving process

It is the process occurs when the solute decomposes or dissociate into negative and positive ions or into separated polar molecules. Each of them binds to the molecules of the solvent.

The mechanism of dissolving process

It's easy to dissolve:

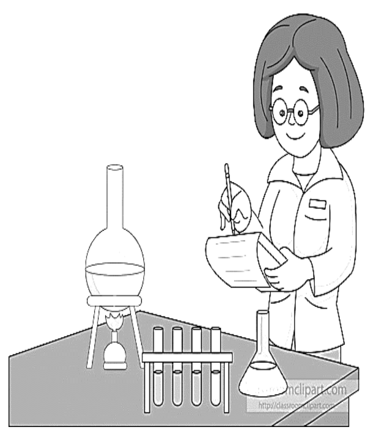
- Ionic compounds (as sodium chloride) **in polar solvent**
- Polar covalent compound (as hydrogen chloride gas) **(as water)**

❖ **The speed of the dissolving process depends on:**

1-Surface area of the solute

2-Stirring

3-Temperature



Solubility

Solubility:

It is the ability of solute to dissolve in a certain amount of solvent.

Or it is the ability of solvent to dissolve certain amount of solute.

Degree of solubility:

It is the mass of solute by grams which dissolve in 100 grams of the solvent to form a saturated solution at standard conditions.

Factors affecting the solubility

1-The nature of solute and solvent:

Like dissolves like

Polar solvent as (water)

Dissolves

Ionic compounds

As NaCl

NaOH

Polar compounds

As HCl

NH_3

Non-polar solvent

Dissolves

Non polar compounds

As oil, fats

-Substances that easily dissolve in **water** are **ionic or polar covalent** compounds.

Give reason

Oil is insoluble in water.

Because oil is non polar while water is polar compound.

Oil is soluble in benzene.

Because both of them is non polar.

Sugar is soluble in water although sugar is non polar.

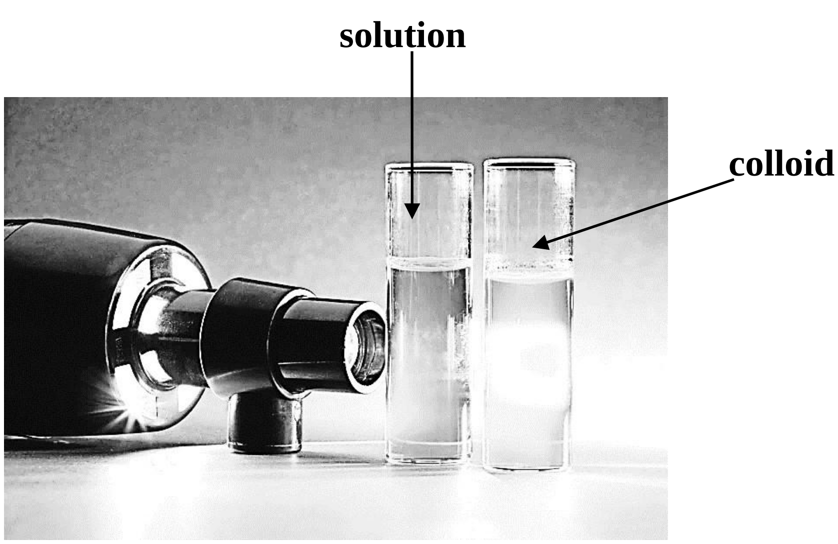
Because sugar molecules make hydrogen bond with water

2-Temperature:

-The solubility of most ionic substance increases with increasing the temperature

Some properties of solution

- Particles cannot be distinguished by naked eye or by the electron microscope.
- Diameter of the particles is less than 1 nm
- Particles of solution are regularly distributed
- Particles don't scatter a beam of light passing through it.



Unit three

Chapter (1): Solution and colloids

Part (2): Concentration of solutions

* You can change the concentration of solution by changing the amount of solute in a solvent.

*The solution will be concentrated if the amount of solute is large (but not larger than solvent).

* The solution is said to be diluted when the amount of solute is small

* We express the concentration of solutions by

(Percentage – molarity – molality)

1-Percentage

Percentage (volume – volume) = $\frac{\textit{solute volume}}{\textit{solution volume}} \times 100$

Percentage (mass – mass) = $\frac{\textit{solute mass}}{\textit{solution mass}} \times 100$

Solution mass = (solute mass + solvent mass)

2-Molarity

It is the number of solute moles that dissolved in one liter of solution

Unit : (Mol / L) or **molar (M)**

Molarity = $\frac{\textit{Number of solute moles (mol)}}{\textit{solution volume (L)}}$

Example: **Calculate the molarity of sugar cane solution $C_{12}H_{22}O_{12}$ in water, if you knew that the mass of the dissolved sugar is 85.5 g in a solution volume of 0.5 L (C = 12, H=1, O=16).**

Answer:

Molar mass of sugarcane = (12 × 12) + (1 × 22) + (12 × 16)=358 g/mol

Number of moles = mass / molar mass = (85.5) / (358) = 0.24 mol

Molarity = $\frac{Number\ of\ moles\ (mol)}{solution\ volume\ (L)} = \frac{0.25}{0.5} = 0.47$ mol/liter

---

3-Molality

It is the number of solute moles in one kilogram of solvent

Unit : (Mol / Kg)

Molality = $\frac{\boldsymbol{Number\ of\ solute\ moles\ (mol)}}{\boldsymbol{solvent\ mass\ (kg)}}$

Example: **Calculate the molality of a prepared solution by dissolving 20 g of sodium hydroxide in 800 g of water, knowing that (Na = 23 , H = 1, O= 16)**

Answer:

Molar mass of sodium hydroxide (NaOH) = (23 + 16 +1) = 40 g /mol

Number of moles = mass / molar mass = 20/40 =0.5 mol.

Mass of solvent by kilogram = 800/1000 = 0.8 kg

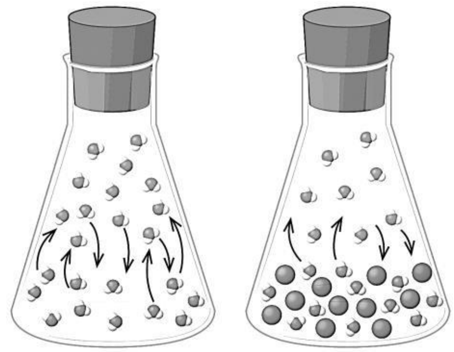
Molality = $\frac{Number\ f\ moles\ (mol)}{solvent\ mass\ (kg)} = \frac{0.5}{0.8} = 0.625$ mol / kg

Unit three

Chapter (1): Solution and colloids

Part (3): Collegative properties of solutions

The properties of a pure solvent differ from its properties by dissolving a solid non volatile substance in it. These properties as (vapor pressure, boiling point, freezing point)



1-Vapor pressure

It is the pressure that exerted by a vapor in dynamic equilibrium with its liquid inside a closed container at a constant temperature and pressure

- **Difference in vapor pressure of pure solvent and solution**

Pure solvent	Solution
-The surface molecules which exposed to vaporization process are the **solvent molecules only**.	-The surface molecules are the **solvent molecules** and **solute molecules** so the number of solvent molecules which exposed to vaporization process **decrease**
-The force that has to be overcome is the attraction force between the solvent molecules with each other.	-The force that has to be overcome is the attraction force between solvent and solute molecule that is larger than the attraction between solvent molecules only.

2-Boiling point:

It is the temperature in which the vapor pressure of the liquid equals the atmospheric pressure.

Measured boiling point:

The temperature at which the vapor pressure of the liquid equals the pressure exerted or acted on it.

- It can be used as indicator for purity of solvent.

- **Pure water boils at 100^0c while salty water causes increasing in the boiling point.(G.R)**

  Because by adding salt to water the vapor pressure of the solution decrease so the solution needs more energy until its vapor pressure equals to atmospheric pressure to boil.

- Boiling point increases by increasing the number of moles of ions in the solution.
- **Boiling point of 0.2M sodium chloride solution is equal to boiling point of 0.2M potassium nitrate solution.(G.R)**

  Because both of them produce the same number of moles of ions in the solution.

- **Boiling point of sodium carbonate is higher than boiling point of sodium chloride with same concentration.(G.R)**

  Because the number of moles of ions in Na_2CO_3 is higher than that of NaCl.

3-Freezing point:

- Freezing point is opposite to boiling point.
- **Freezing point of solution is <u>less than</u> freezing point of pure solvent.(G.R)**

Because the attraction force between solvent and solute increase so number of solvent molecules that will change into solid state on freezing decrease.

- Decreasing in Freezing point is inversely proportional to the number of dissolved solute in the solutions.
- Freezing point of sugary solution (not ionized into ions) is -1.86 ^{0}C.
- Freezing point of sodium chloride (produce two ions) is

  (2 x -1.86) = - 3.72^{0}C.

- **Salt is added to snow – covered roads in cold places.(G.R)**

Because salt decrease the freezing point of water so water will not easily change to snow and this prevent cars from skidding and decrease the number of accidents.

Suspension

They are heterogeneous solutions in which you can distinguish its components by your eye.

- The suspended particles precipitate if it lift for a short time without shaking.
- The diameter of its particles is larger than 1000 nm.
- The suspended particles can be seen by eye.
- The suspended particles can be separated by filtration as filter paper hold the suspended particles while water pass through paper.
- **<u>Examples</u> (Sand in water – Chalk powder in water)**

Colloids

They are heterogeneous mixtures that carry the properties of solution and suspension.

- The dispersed particles don't precipitate if they are left for a short time without shaking.
- The dispersed particles diameter is from 1- 1000 nm
- The dispersed particles can be seen by electronic microscope only
- The dispersed particles cannot be separated by filtration.
- The shape depends on its concentration

- **Concentrated colloids appears as milk**
- **Diluted colloids appears clear**

Dispersed phase	Dispersed Medium	Examples
Gas	Liquid	Some types of creams – whipped egg
	Solid	Sweat made of sugar and egg white
Liquid	Gas	Aerosols
	Liquid	Milk – mayonnaise
	Solid	Hair gel
Solid	Gas	Dust in air particles
	Liquid	Pigment – blood – starch in hot water

- **There is no gas-gas colloidal system.(G.R)**

  Because mixed gases are homogenous mixture

Preparation methods for colloids

1-Dispersion method:

1- The substance is crushed into small particles until its size reaches the size of colloid particles.

2- Then added to the dispersed medium with stirring

<u>As</u> (Starch in hot water)

2-Condensation method:

The small particles are collected together into larger particles have the volume of the colloid particles by chemical reactions as

(oxidation – reduction – hydrolysis)

Oxidation

$$2H_2S_{(aq)} + SO_{2(g)} \longrightarrow 3S_{(colloid)} + 2H_2O_{(l)}$$

Reduction

Unit three
Chapter (2): Acids and bases

Part (1): Properties of Acids and Bases

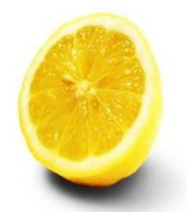
❖ **Industries including acids:**

1-Fertilizers 2-Medicines 3-Plastic 4-Car batteries

❖ **Industries including bases**

1-Soap 2-Detergents 3-Dyes 4-Medicines



***Natural and artificial products including acids or bases in their composition:**

Product	Acids entering in its composition
Acidic plants (lemon – oranges – tomatoes)	Citric acid – Ascorbic acid
Dairy products (Milk – yoghurt)	Lactic acid
Soft drinks	Carbonic acid – phosphoric acid

Product	Bases entering in its composition
Soap	Sodium hydroxide
Baking soda	Sodium bicarbonate
Washing soda	Hydrated sodium carbonate

*Properties of acids and bases

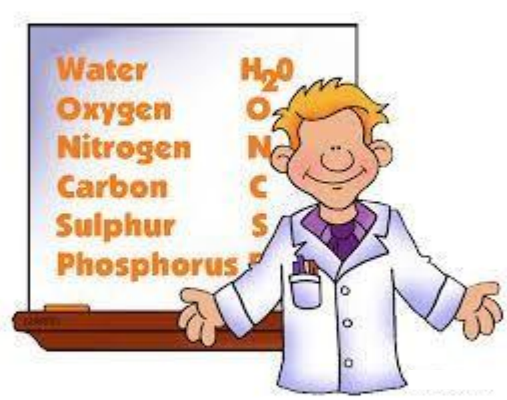
p.o.c	Acids	Bases
Taste	Sour taste	Bitter taste
Effect on litmus paper	Change the color of litmus into red	Change the color of litmus into blue
Reactions	* With active metals to give salt of acid and hydrogen gas $Zn + 2HCl \longrightarrow ZnCl_2 + H_2$ *With carbonate and bicarbonate to produce CO_2 $Na_2CO_3 + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O + CO_2$ *With bases to produce salt and water.	*With acids to produce salt and water. $NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$

Theories that describe acid and base

1) Arrhenius theory

2) Bronsted lowery theory

3) Lewis theory



1) Arrhenius theory

1)Definition of acid	**It is the substance that ionize or dissociate in water to give one or more hydrogen ions H^+** $HCl_{(g)} \xrightarrow{Water} H^+_{(aq)} + Cl^-_{(aq)}$ $H_2SO_{4\,(aq)} \xrightarrow{Water} H^+_{(aq)} + HSO_4^-{}_{(aq)}$ $HClO_{4(aq)} \xrightarrow{Water} H^+_{(aq)} + ClO_4^-{}_{(aq)}$ So acid is good conductor of electricity. And it increases the concentration of positive hydrogen ions in aquatic solutions.
2)Definition of base	**It is the substance that ionize or dissociate into water to give one or more hydroxide ions OH^-** $NaOH_{(S)} \xrightarrow{water} Na^+_{(aq)} + OH^-_{(aq)}$ $KOH_{(aq)} \xrightarrow{water} K^+_{(aq)} + OH^-_{(aq)}$ $Ba(OH)_{2\,(aq)} \xrightarrow{water} Ba^{+2}_{(aq)} + 2OH^-_{(aq)}$ So base is good conductor of electricity. And it increases the concentration of negative hydroxide ions in solution.
3)Reaction between acid an base	*It produces salt and water. $HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$ • The neutralization reaction $H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$
4)Observations on Arrhenius theory	1) Carbon dioxide doesn't contain a source of positive hydrogen ion but is considered as acid. 2) He said that acid must contain hydrogen ion and base must contain hydroxide group and it is not completely correct. **Ammonia** • In water give hydroxide ion while it is not Arrhenius base. $NH_3 + H_2O \longrightarrow NH_3^+ + OH^-$ It neutralizes with acid

2) Bronsted Lowry theory

1)Definition of acid	**It is the substance that give the proton H^+ (proton donor).**
2)Definition of base	**I t is the substance that has the ability to accept the proton (proton acceptor).**
3)reaction between acid and base	$\mathbf{HA} + \mathbf{B} \longrightarrow \mathbf{A^-} + \mathbf{BH^+}$ Acid Base conjugate base Conjugate acid
4)Examples	**1) Hydrogen chloride and water** $HCl + H_2O \longrightarrow Cl^- + H_3O^+$ (acid) (base) conjugate base conjugate acid **2)Ammonia in water** $NH_3 + H_2O \longrightarrow OH^- + NH_4^+$ (base) (Acid) conjugate base conjugate acid
5)Notes	**<u>Conjugate acid</u>** The substance that produced when base accepts a proton. **<u>Conjugate base</u>** The substance that produced when acid loses a proton.

3) Lewis theory

1)Definition of acid	Substance that accept an electron pair or more
2)Definition of base	Substance that donates an electron pair or more
3)Examples	**Reaction of hydrogen ion with fluoride ion** $H^+ \quad + \quad F^- \longrightarrow HF$ Lewis acid Lewis base

Comparison of acid and base in the three theories

Theory	Acid definition	Base definition
Arrhenius	H+ producer	OH- producer
Bronsted – Lowry	H+ donor	H+ acceptor
Lewis	Electron pair acceptor	Electron pair donor

Unit three

Chapter (2): Acids and bases

Part (2): Classification of acids and bases

1) According to its source into:

Organic acids

*Acids that have an organic origin (Plant or animal)

*All of them are weak acids

Examples

-Lactic acid

-Acetic acid

-Citric acid

-Oxalic acid

-Formic acid

Mineral acids

*Acids that have no organic origin

Or have non metallic element in their structure.

*Some are weak and some are strong

Examples

-Carbonic acid

-Hydrochloric acid

-Phosphoric acid

-Sulphuric acid


classroomclipart.com

2) According to the number of hydrogen atoms that the acid react through it (basisty of acid)

1)Mono basic acids	2)Dibasic acids	3)Tribasic acids
When it dissolves in water each molecule gives one proton.	When it dissolves in water each molecule gives one or two protons.	They are acids that can give three protons through reactions.
Examples: -Hydrochloric acid (HCl) -Nitric acid (HNO_3) -Acetic acid (CH_3COOH) -Formic acid($HCOOH$)	**Examples:** -Sulphuric acid (H_2SO_4) -Carbonic acid (H_2CO_3) -Oxalic acid COOH | COOH	**Examples:** -Phosphoric acid (H_3PO_4) -Citric acid CH_2— COOH | H O——C—— COOH | CH_2—COOH

3) According to its strength into:

1)Strong acids	2)Weak acids
Acids which are completely ionized in water	Acids which are incompletely ionized in water
Examples: -Hydrochloric acid (HCl) -Nitric acid (HNO_3) -Sulphuric acid (H_2SO_4)	**Examples:** -Acetic acid (CH_3COOH) -Formic acid($HCOOH$) -Oxalic acid COOH | COOH

Classification of bases

1) According to its molecular composition

Base	Examples	Application
1-Metal oxides	Iron (II)oxide FeO	$FeO + 2\ HCl \longrightarrow FeCl_2 + H_2O$
2-Metal hydroxide	Calcium hydroxide $Ca(OH)_2$	$Ca(OH)_2 + H_2SO_4 \longrightarrow CaSO_4 + 2\ H_2O$
3-Metal carbonate	Potassium carbonate K_2CO_3	$K_2CO_3 + 2HCl \longrightarrow 2KCl + H_2O + CO_2$
4-Metal bicarbonates	Potassium bicarbonate $KHCO_3$	$KHCO_3 + HCl \longrightarrow KCl + H_2O + CO_2$

2) According to its strength:

1)Strong bases	2)Weak bases
Bases which are completely ionized in water	Bases which are incompletely ionized in water
Examples: -Potassium hydroxide (KOH) -Sodium hydroxide (NaOH)	**Examples:** -Ammonium hydroxide (NH_4OH)

***Bases that dissolve in water are called alkalis.**

So all alkalis are bases but not all bases are alkalis.

Detecting acids and bases

-By PH meter or indicators

1) Indicators:

They are weak organic acids or bases their color changes with the change of the solution type.

Indicator	Colour in acidic medium	Colour in neutral medium	Colour in basic medium
Methyl orange	Red	Orange	Yellow
Bromothymol blue	Yellow	Green	Blue
Phenolphthalein	Colourless	Colourless	Pink
Litmus	Red	Violet	Blue

2)By PH meter

PH < 7 so the substance is acid

PH = 7 so the substance is neutral

PH > 7 so the substance is basic

Unit three

Chapter (2): Acids and bases

*Salt are found in earth crust, dissolved in sea water or precipitated in the seabed

<u>*Salt formed from</u>

a) Cation: A positive ion of the base.

b) Anion: A negative ion of the acid.

Example: potassium nitrate

K^+ NO_3^-

Cation Anion

Notes

<u>1-Monobasic acids form only one type of salts</u>

As (Nitric acid HNO_3)..................forms nitrates salts only

<u>2-Dibasic acids form two type of salts</u>

As (Sulphuric acid H_2SO_4)........ forms sulphate and bisulphate salts

<u>3-Tribasic acids form three type of salts</u>

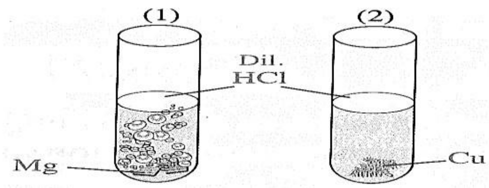
As (Phosphoric acid H_3PO_4) ....forms three types of salts

Formation of salts

1) Reaction of diluted acid with active metals:

When metal is more active than hydrogen

$$Zn_{(s)} + H_2SO_{4(aq)} \xrightarrow{\text{dil.}} ZnSO_{4\,(aq)} + H_{2(g)}\uparrow$$



2) Reaction of metal oxides with acids

When metal is less active than hydrogen

$$CuO_{(S)} + H_2SO_{4(aq)} \xrightarrow{\triangle} CuSO_{4(aq)} + H_2O_{(l)}$$

3) Reaction of metal hydroxides with acids:

$$NaOH_{(aq)} + HCl_{(aq)} \xrightarrow{\triangle} NaCl_{(aq)} + H_2O_{(l)}$$

4) Reaction of metal carbonates or bicarbonates with acid (acidity test):

$$Na_2CO_{3(aq)} + 2HCl_{(aq)} \longrightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_2\uparrow$$

Types of aqueous solution of salts

1-Neutral salts:

Produced from reaction of

a) Strong acid and strong base

as: NaCl produced from (NaOH & HCl)

b) Weak acid and weak base

As: CH_3COONH_4 produced from (CH_3COOH & NH_4OH)

2-Acidic salts:

Produced from reaction of

Strong acid and weak base

As: NH_4Cl produced from (NH_4OH & HCl)

3-Basic salts:

Produced from reaction of

Strong base and weak acid

As: Na_2CO_3 produced from (NaOH & H_2CO_3